Name ____ AP Chemistry

Determining the Molar Volume of a Gas

Introduction

From blimps to airbags, gases are used to fill a wide variety of containers. How much of a particular gas must be produced to fill a container? The amount of gas needed to fill any size container can be calculated if the molar volume of a gas is known.

Avogadro's law states that equal volumes of gases contain an equal number of molecules under the same conditions of temperature and pressure. It follows, therefore, that all gas samples containing the same number of molecules will occupy the same volume if the temperature and pressure are kept constant. The volume occupied by one mole of a gas is called molar volume. In this experiment, the molar volume of hydrogen gas at standard temperature and pressure (STP, equal to 273 K and 1 atm) will be measured.

The reaction of magnesium metal with hydrochloric acid (Equation 1) provides a convenient means of generating hydrogen in the lab:

$$Mg\ (s) + 2\ HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2\ (g) \qquad \text{Equation 1}$$

If the reaction is carried out with excess hydrochloric acid, the volume of hydrogen gas obtained will depend on the number of moles of magnesium as well as the pressure and temperature. The molar volume of hydrogen can be calculated if the volume occupied by a sample containing a known number of moles of hydrogen is measured. Since the volume will be measured under laboratory conditions of temperature and pressure, the measured volume must be corrected to STP conditions before calculating the molar volume.

The relationship among the four gas variables—pressure (P), volume (V), temperature (T), and the number of moles (n)—is expressed in the ideal gas law (Equation 2), where R is a constant called the universal gas constant:

$$PV = nRT \qquad \text{Equation 2}$$

The ideal gas law reduces to Equation 3, the combined gas law, if the number of moles of gas is constant. The combined gas law can be used to calculate the volume (V_2) of a gas at STP (T_2 and P_2) from the volume (V_1) measured under any other set of laboratory conditions (T_1 and P_1). In using either the ideal gas law or the combined gas law, remember that temperature must be always expressed in units of kelvins (K) on the absolute temperature scale:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \text{Equation 3}$$

Hydrogen gas will be collected by the displacement of water in an inverted burette, as shown in Figure 1. The total pressure of the gas in the tube will be equal to the barometric (air) pressure. However, the gas in the burette will not be pure hydrogen. The gas will also contain water vapor due to the evaporation of the water molecules over which hydrogen is being collected. According to Dalton's law, the total pressure of the gas will be equal to the partial pressure of hydrogen plus the partial pressure of water vapor (Equation 4). The vapor pressure of water depends only on the temperature (see Table 1).

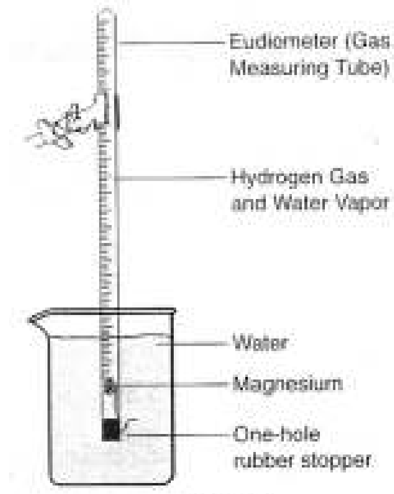

Figure 1.

$$P_{total} = P_{H_2} + P_{H_2O} \quad \text{Equation 4}$$

Temperature (°C)	Pressure (H_2O) (mm Hg)	Temperature (°C)	Pressure (H_2O) (mm Hg)
16	13.6	22	19.8
17	14.5	23	21.1
18	15.5	24	22.4
19	16.5	25	23.8
20	17.5	26	25.2
21	18.7	27	26.7

The purpose of this experiment is to determine the volume of one mole of hydrogen gas at standard temperature and pressure (STP). Hydrogen will be generated by the reaction of a known mass of magnesium with excess hydrochloric acid in an inverted burette filled with water. The volume of hydrogen collected by water displacement will be measured and corrected for differences in temperature and pressure in order to calculate the molar volume of hydrogen at STP.

Prelab Questions

A reaction of 0.028 g of magnesium with excess hydrochloric acid generated 31.0 mL of gas. The gas was collected by water displacement in a 22°C water bath. The barometric pressure in the lab that day was 746 mm Hg.

1) Use Dalton's law and the vapor pressure of water at 22°C (Table 1) to calculate the partial pressure of hydrogen gas in the burette.
2) Use the combined gas law to calculate the "corrected" volume of hydrogen at STP (hint: watch the units for temperature and units).

3) What is the theoretical number of moles of hydrogen that can be produced from 0.028 g of Mg? (Note: refer to Equation 1 for the balanced equation of the reaction.)
4) Divide the corrected volume of hydrogen by the theoretical number of moles of hydrogen to calculate the molar volume (in L/mol) of hydrogen at STP.

Procedure

1) Fill a 400 mL beaker about ¾ full of water.
2) Obtain or cut a 4.5 cm piece of magnesium ribbon. Measure and record in the data table the exact length of the magnesium ribbon to the nearest 0.1 cm.
3) Using the density of magnesium of 0.0094 g/cm, determine the mass of the magnesium. Record this mass in the data table.
4) Fold the magnesium ribbon into a small ball.
5) Obtain a piece of copper wire about 15 cm long. Twist one end of the copper wire around the ball of magnesium ribbon, creating a cage of copper wire around the magnesium ribbon.
6) Obtain 15 mL of 2 M hydrochloric acid in a 25 mL graduated cylinder.
7) Obtain a burette and ensure that the stopcock at the end is closed. Tipping the burette to the side, slowly add the 2 M hydrochloric acid.
8) Move the burette over a sink, keeping it in the tipped position. Use a wash bottle to slowly and carefully fill the burette with water. Work slowly to avoid mixing the acid and water layers. Fill the burette all the way to the top so no air remains in the tube.
9) Obtain a rubber stopper with a hole and feed the copper wire through the hole until the cage of magnesium ribbon is approximately 7 cm below the bottom of the stopper. Hook the end of the copper wire around the top of the stopper to hold the cage of magnesium ribbon in place.
10) Insert the rubber stopper with the copper wire into the burette.
11) Place the 400 mL beaker filled with water on a ring stand with a clamp.
12) Place your finger over the hole of the rubber stopper. Quickly invert the burette and place it into the 400 mL, clamping the burette into place.
13) The hydrochloric acid (which is denser than water) should diffuse downward through the burette and begin to react with magnesium. Allow the apparatus to sit for 5 minutes after the magnesium has completely reacted. Gently tap the sides of burette to dislodge any gas bubbles that may have become attached to the sides.
14) Fill a 500 mL graduated cylinder with tap water and place the cylinder in a sink.
15) Carefully unclamp the burette. Sticking your hand under water, place your finger over the hole in the rubber stopper. Transfer the burette to the 500 mL graduated cylinder.
16) Move the burette up and down until the water levels on the inside and outside of the burette is equal (note: make sure to keep the rubber stopper beneath water). Measure and record the volume in the data table.

17) Measure and record the temperature of the water within the graduated cylinder in the data table.
18) Record the barometric pressure in the data table.
19) Remove the burette from the graduated cylinder. Discard the solution in the sink and clean out the burette.
20) Repeat the procedure for a second trial if time permits.

Data Table

	Trial #1	**Trial #2**
Length of Mg		
Mass of Mg		
Volume of H_2 Gas		
Barometric Pressure		

Postlab Questions

1) Calculate the theoretical number of moles of hydrogen gas produced in Trials 1 and 2.
2) Use Table 1 to find the vapor pressure of water at the temperature of the water in this experiment. Calculate the partial pressure of hydrogen gas produced in Trials 1 and 2.
3) Use the combined gas law to convert the measured volume of hydrogen to the "ideal" volume the hydrogen gas would occupy at STP for Trials 1 and 2.
4) Divide the volume of hydrogen gas at STP by the theoretical number of moles of hydrogen to calculate molar volume of hydrogen for Trials 1 and 2.
5) What is the average value of the molar volume of hydrogen? Using the definition value of molar value, calculate the percent error in your experimental determination of the molar volume of hydrogen.
6) One mole of hydrogen has a mass of 2.02 g. Use your value of molar volume of hydrogen to calculate the mass of one litter of hydrogen gas at STP. This is the density of hydrogen in g/L. Given that the density of hydrogen is 0.0899 g/L, how does this experimental value of the density compare?
7) In setting up the experiment, a student noticed that a bubble of air leaked into the burette when it was inverted in the water. What effect would this have on the measured volume of hydrogen gas? Would the calculated molar volume be too high or too low as a result of this error? Explain.
8) A student noticed that the magnesium ribbon appeared oxidized—the metal surface was black and dull rather than silver and shiny. What effect would this error have on the measured volume of hydrogen gas? Would the calculated molar volume of hydrogen be too high or too low as a result of this error? Explain.